Quick Problems Using Equations in Chapter 10 and 11

*Clausius-Clapeyron Equation* $\ln\left(\frac{P_{vap_1}}{P_{vap_2}}\right) = \frac{\Delta H}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$

If the barometric pressure in Reno, NV (elevation 4500 ft) is 680 torr, at what temperature will water boil?

$P_{vap_1} = 760\text{ torr}$ $T_2 = ?$

$P_{vap_2} = 680\text{ torr}$ $T_1 = 100.°C$ $373$

$\Delta H = 40.7\frac{kJ}{mol}$

$$\ln\left(\frac{760\text{ torr}}{680\text{ torr}}\right) = \frac{40.7\frac{kJ}{mol}}{8.314\times10^{-3}\frac{kJ}{mol\,k}}\left(\frac{1}{T_2} - \frac{1}{373k}\right) \quad T_2 = 370k - 273 = \boxed{96.9°C}$$

The melting point of potassium metal is 63.2°C. Molten potassium has a vapor pressure of 10.00 torr at 443°C and a vapor pressure of 400.0 torr at 708°C. Use these data to calculate the heat of vaporization of liquid potassium.

$$\ln\left(\frac{400.0\text{ torr}}{10.00\text{ torr}}\right) = \frac{\Delta H}{8.314\frac{J}{mol\,k}}\left(\frac{1}{716k} - \frac{1}{981k}\right) \rightarrow \frac{\cancel{4176\,J}}{mol}\ 81300 = \cancel{118}\,\frac{kJ}{mol}\ 81.3\frac{kJ}{mol}$$

*Heating Curves*

Calculate the enthalpy change upon converting 50.0 g of ice at -25.0°C to water vapor at 125°C under a constant pressure of 1 atm. The specific heats of ice and steam are 2.09 J g$^{-1}$ °C$^{-1}$, and 1.84 J g$^{-1}$ °C$^{-1}$ respectively. For $H_2O$, $\Delta H_{fus} = 6.01$ kJ mol$^{-1}$ and $\Delta H_{vap} = 40.7$ kJ mol$^{-1}$.

$\times \frac{1 mol}{18.0 g} = 2.77$

$\Delta H_1 = (50.0g)(2.09\frac{J}{g°C})(25°C) \times \frac{1 kJ}{1000 J} = 2.61\,kJ$

$\Delta H_2 = 2.78\,mol\,H_2O \times \frac{6.01\,kJ}{mol} = 16.7\,kJ$

$\Delta H_3\ 50.0g(4.18\frac{J}{g°C})(100°C) \times \frac{1kJ}{1000J} = 20.9\,kJ$

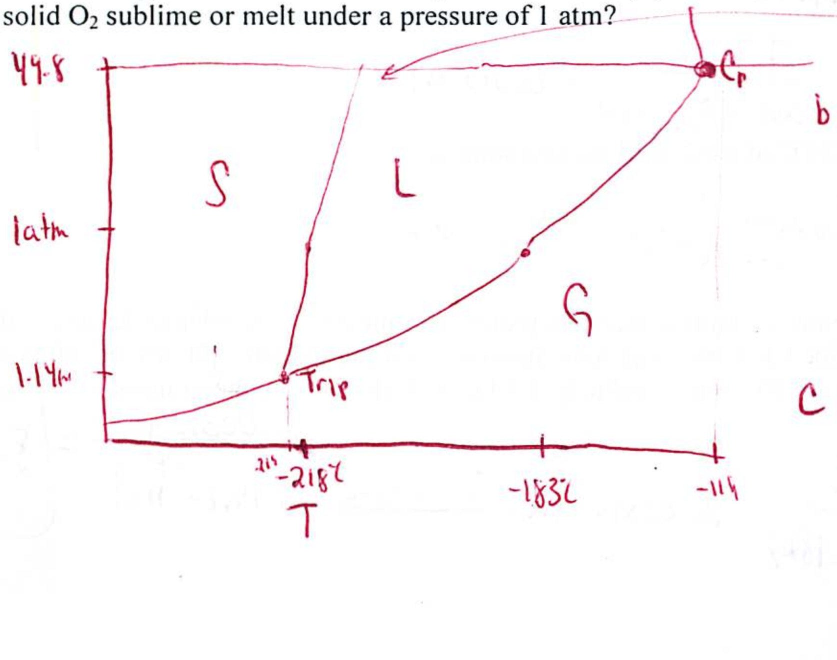
$\Delta H_4\ 2.78\,mol\,H_2O \times \frac{40.7\,kJ}{mol} = 113\,kJ$

$\Delta H_5\ 50.0g(1.84\frac{J}{g°C})(25°C)\frac{1kJ}{1000J} = 2.3\,kJ$

$\boxed{\Delta H = 156\,kJ}$

*Phase Diagrams*

The normal melting and boiling points of $O_2$ are -218°C and -183°C, respectively. Its triple point is at -219°C and 1.14 torr, and its critical point is at -119°C and 49.8 atm. (a) Sketch the phase diagram for $O_2$, showing the four points given and indicating the area in which each phase is stable. (b) Will $O_{2(s)}$ float on $O_{2(l)}$? Explain. (c) As it is heated, will solid $O_2$ sublime or melt under a pressure of 1 atm?




b No B/c the line shows that @ higher P $O_2$ Freezes There fore the molecule are closely packed in solid there fore less denser.

c melt b/c you will passing through the S-L line.

*Henry's Law* ~~[illegible]~~ $C = kP$

Calculate the concentration of $CO_2$ in a soft drink that is bottled with a partial pressure of $CO_2$ of 4.0 atm over the liquid at 25°C. The Henry's law constant for $CO_2$ in water at this temperature is $3.1\text{x}10^{-2}$ mol $L^{-1}$ $atm^{-1}$. Calculate the concentration of $CO_2$ in a soft drink after the bottle is opened and equilibrates at 25°C under a partial pressure of $3.0\text{x}10^{-4}$ atm.

$$C = \left(3.1\times10^{-2}\,\frac{M}{atm}\right)(4.0\,atm) = 0.12\,M$$

$$C\left(3.1\times10^{-2}\,\frac{M}{atm}\right)(3.0\times10^{-4}\,atm) = 9.3\times10^{-6}\,M$$

*Raoult's Law* $P_{vap} = \chi_{H_2O} P^\circ_{vap}$

Glycerin ($C_3H_8O_3$) is a nonvolatile nonelectrolyte with a density of 1.26 g/mL at 25°C. Calculate the vapor pressure at 25°C of a solution made by adding 50.0 mL of glycerin to about 500.0 mL of water. The vapor pressure of pure water at 25°C is 23.8 torr.

$$50.0\,mL \times \frac{1.26\,g}{mL} = 63\,g \times \frac{1\,mol\,Gly}{76\,g\,Gly} = 0.829\,mol\,Gly$$

$$500.0\,mL \times \frac{1.0\,g}{mL} \times \frac{1\,mol\,H_2O}{18\,g\,H_2O} = 27.8\,mol\,H_2O$$

$$\chi_{H_2O} = \frac{27.8\,mol}{27.8\,mol + 0.829\,mol} = 0.972\,mol$$

$$P_{vap} = (0.972\,mol)(23.8\,torr) = \boxed{23.1\,torr}$$

Calculate the mass of ethylene glycol ($C_2H_6O_2$) that must be added to 1.00 kg of ethanol ($C_2H_5OH$) to reduce its vapor pressure by 10.0 torr at 35°C. The vapor pressure of pure ethanol at 35°C is $1.00\text{x}10^2$ torr.

$$mol\,EtOH = 1.00\,kg\,EtOH \times \frac{1000\,g}{1\,kg} \times \frac{1\,mol\,EtOH}{46\,g\,C_2H_5OH} = 21.7\,mol$$

$$90\,torr = \chi_{EtOH}\ 100\,torr$$

$$\chi_{EtOH} = 0.9 = \frac{mol\,EtOH}{mol\,EtOH + mol\,EtGly} \Rightarrow 0.9 = \frac{21.7\,mol}{21.7\,mol + x} \Rightarrow 21.7\,mol + x = 24.1\,mol$$

$$x = 2.4\,mol\,EtGly \times \frac{62\,g\,EtGly}{1\,mol\,EtGly} = \boxed{149\,g}$$

*Osmotic Pressure*

The average osmotic pressure of blood is 7.7 atm at 25°C. What concentration of glucose in aqueous solution will be isotonic with blood? $\Pi = MRT$

$$M = \frac{\Pi}{RT} = \frac{7.7\,atm}{0.0821\,\frac{L\,atm}{mol\,K}\ 298\,K} = 0.315\,M$$

What is the osmotic pressure at 20°C of a 0.200 M sucrose solution?

$$(0.200\,M)\left(0.0821\,\frac{L\,atm}{mol\,K}\right)(293\,K) = 4.81\,atm$$

The osmotic pressure of an aqueous solution of a certain protein was measured in order to determine the protein's molar mass. The solution contained 3.50 mg of protein dissolved in water to form 5.00 mL of solution. The osmotic pressure of the solution at 25°C was found to be 1.54 torr. Calculate the molar mass of the protein.

$$1.54\,torr \times \frac{1\,atm}{760\,torr} = \frac{0.00202\,atm}{\left(0.0821\,\frac{L\,atm}{mol\,K}\right)(298\,K)} = 8.28\times10^{-5}\,M \times 0.00500\,L = 4.14\times10^{-7}\,mol$$

$$\frac{0.00350\,g}{4.14\times10^{-7}\,mol} = \boxed{8.45\times10^{3}\,\frac{g}{mol}}$$